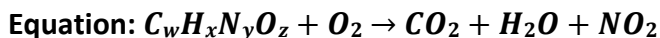


1. The combustion of 1.38 g of a compound which contains C, H, O, and N yields 1.72 g of CO₂ and 1.18 g of H₂O. Another sample of the compound with a mass of 22.34 g is found to contain 6.75 g of O. What is the empirical formula of the compound?



When asked for empirical formula, think about what information you need to eventually take mole ratios. This problem is very similar to yesterday's problem, just with another element to deal with. The approach however, is essentially the same.

$$1.18 \text{ g H}_2\text{O} * \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} * \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 0.13097 \text{ mol H} * \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 0.1320 \text{ g H}$$

$$1.72 \text{ g CO}_2 * \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} * \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 0.03908 \text{ mol C} * \frac{12.011 \text{ g C}}{1 \text{ mol C}} = 0.4694 \text{ g C}$$

$$\text{Mass \% of O} = \frac{6.75 \text{ g}}{22.34 \text{ g}} * 100\% = 30.21\%$$

Now, let's find the amount of O in the sample that was combusted.

$$\text{Mass O} = \text{Mass \% O} * \text{sample mass} = 0.3021 * 1.38 \text{ g} = 0.4170 \text{ g O}$$

Find the moles of O:

$$0.4170 \text{ g O} * \frac{1 \text{ mol O}}{15.9994 \text{ g O}} = 0.02606 \text{ mol O}$$

Use the calculated masses to find the mass of N in the sample, and subsequently moles of N.

$$\begin{aligned} \text{Mass N} &= \text{Mass sample} - \text{Mass C} - \text{Mass H} - \text{Mass O} \\ &= 1.38 \text{ g} - 0.4694 \text{ g} - 0.1320 \text{ g} - 0.4170 \text{ g} = 0.3616 \text{ g N} \end{aligned}$$

$$0.3616 \text{ g N} * \frac{1 \text{ mol N}}{14.0067 \text{ g N}} = 0.02582 \text{ mol N}$$

Take mole ratios to find empirical formula!

$$\text{Mass \% of O} = \frac{6.75 \text{ g}}{22.34 \text{ g}} * 100\% = 30.21\%$$

Now, let's find the amount of O in the sample that was combusted.

$$\text{Mass O} = \text{Mass \% O} * \text{sample mass} = 0.3021 * 1.38 \text{ g} = 0.4170 \text{ g O}$$

$$\text{C: } \frac{0.03908}{0.02582} = 1.5$$

$$\text{H: } \frac{0.13097}{0.02582} = 5.07 \sim 5$$

$$\text{O: } \frac{0.02606}{0.02582} = 1$$

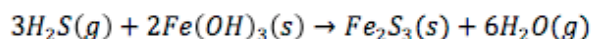
$$\text{N: } \frac{0.02582}{0.02582} = 1$$

Resulting formula: $\text{C}_{1.5}\text{H}_5\text{NO}$

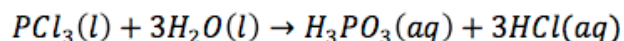
Remember empirical formulas must contain whole number coefficients! The final answer is: $\text{C}_3\text{H}_{10}\text{N}_2\text{O}_2$

2. Write balanced equations corresponding to the following descriptions.

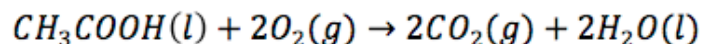
- a. When hydrogen sulfide gas is passed over solid hot iron(III) hydroxide, the resultant reaction produces solid iron(III) sulfide and gaseous water.



- b. When liquid phosphorus trichloride is added to water, it reacts to form aqueous phosphorous acid, $\text{H}_3\text{PO}_3(aq)$, and aqueous hydrochloric acid.



- c. The complete combustion of acetic acid (CH_3COOH), the main active ingredient in vinegar.



3. Washing soda, a compound used to prepare hard water for washing laundry. Its formula is represented as $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$, where x is the number of moles of H_2O per mole of Na_2CO_3 . When a 2.558 g sample of washing soda is heated, all of the water of hydration is lost, leaving 0.948 g of anhydrous Na_2CO_3 left. What is x ?

$$\begin{aligned}\text{Mass } \text{H}_2\text{O} &= \text{Mass hydrate} - \text{Mass anhydrate} \\ &= 2.558 \text{ g } \text{Na}_2\text{CO}_3 \cdot \text{H}_2\text{O} - 0.948 \text{ g } \text{Na}_2\text{CO}_3 = 1.61 \text{ g } \text{H}_2\text{O}\end{aligned}$$

$$1.61 \text{ g } \text{H}_2\text{O} * \frac{1 \text{ mol } \text{H}_2\text{O}}{18.02 \text{ g } \text{H}_2\text{O}} = 0.0894 \text{ mol } \text{H}_2\text{O}$$

$$0.948 \text{ g } \text{Na}_2\text{CO}_3 * \frac{1 \text{ mol } \text{Na}_2\text{CO}_3}{105.988 \text{ g } \text{Na}_2\text{CO}_3} = 0.00894 \text{ mol } \text{Na}_2\text{CO}_3$$

Take mole ratios! Just like finding empirical formula.

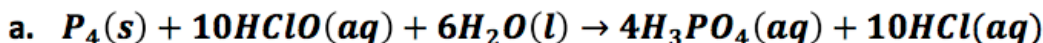
$$\frac{\text{mol } \text{H}_2\text{O}}{\text{mol } \text{Na}_2\text{CO}_3} = \frac{0.0894}{0.00894} = 10$$

For every mole of Na_2CO_3 , there are 10 moles of H_2O . Thus, $x = 10$.

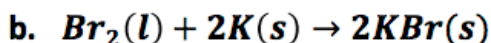
4. Determine the oxidation number of each element in each of the following substances:

a. SO_2	S: +4; O: -2
b. COCl_2	C: +4; O: -2; Cl: -1
c. HBrO	H: +1; Br: +1; O: -2
d. BaCrO_4	Ba: +2; Cr: +6; O: -2
e. HClO_4	H: +1; Cl: +7; O: -2

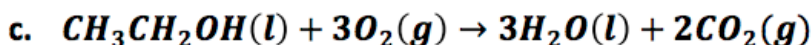
5. Which of the following are redox reactions? For those that are, indicate which elements are being oxidized and reduced. For those that are not, indicate whether they are neutralization or precipitation reactions.



Redox reaction. P is oxidized (0 to +5) and Cl is reduced (+1 to -1)

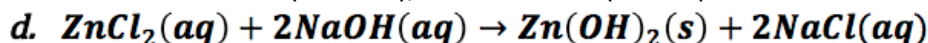


Redox reaction. K is oxidized (0 to +1) and Br is reduced (0 to -1)



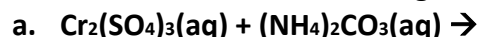
It may help to rewrite the first reactant as $\text{C}_2\text{H}_6\text{O}$. Redox reaction.

C is oxidized (-2 to +4), O_2 is reduced (0 to -2)

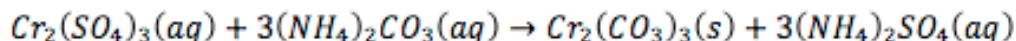


Precipitation reaction (check ox #s to make sure they don't change from reactants to products)

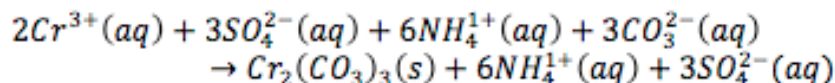
6. Write the overall balanced equation, total ionic equation, and net ionic equations for each of the following cases. Identify the spectator ions.



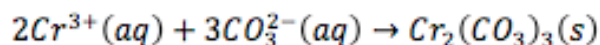
Overall balanced equation:



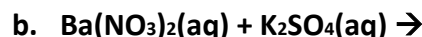
Total ionic equation:



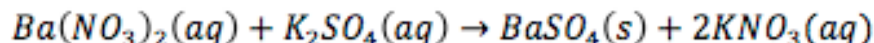
Net ionic equation:



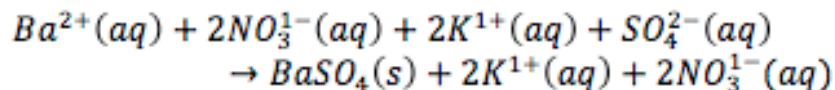
Spectator ions: SO_4^{2-} , NH_4^{1+}



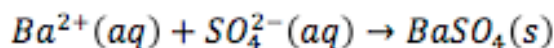
Overall balanced equation:



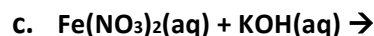
Total ionic equation:



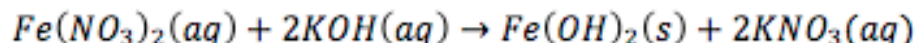
Net ionic equation:



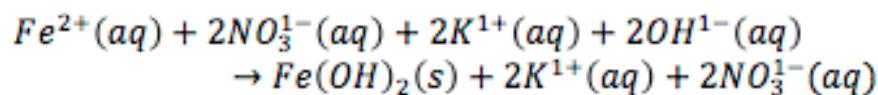
Spectator ions: NO_3^{1-} , K^{1+}



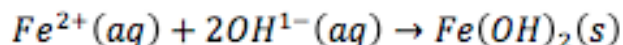
Overall balanced equation:



Total ionic equation:



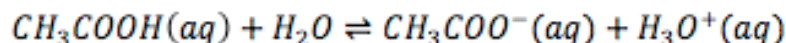
Net ionic equation:



Spectator ions: NO_3^- , K^+

7. Write the balanced chemical and net ionic equations for the following situations.

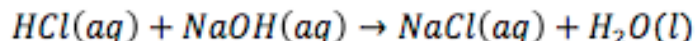
- a. Reaction of acetic acid (CH_3COOH) with water



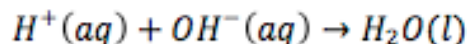
The equation above is the overall equation and the net ionic one since acetic acid is a weak acid.

- b. $HCl(aq) + NaOH(aq) \rightarrow$

Overall equation:

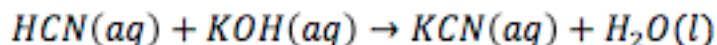


Net ionic equation:

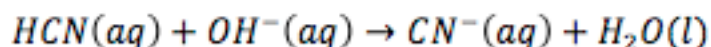


c. $\text{HCN}(\text{aq}) + \text{KOH}(\text{aq}) \rightarrow$ (hint: HCN is a weak acid)

Overall equation:

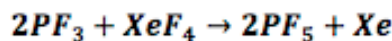


Net ionic equation:



8.

a. How many grams of PF_5 can be formed from 9.46 g of PF_3 and 9.42 g of XeF_4 in the following unbalanced reaction.



Make sure to balance the chemical equation!

$$9.46 \text{ g PF}_3 * \frac{1 \text{ mol PF}_3}{87.97 \text{ g PF}_3} * \frac{2 \text{ mol PF}_5}{2 \text{ mol PF}_3} = 0.108 \text{ mol PF}_5$$

$$9.42 \text{ g XeF}_4 * \frac{1 \text{ mol XeF}_4}{207.28 \text{ g XeF}_4} * \frac{2 \text{ mol PF}_5}{1 \text{ mol XeF}_4} = 0.0909 \text{ mol PF}_5$$

Since 0.0909 moles is less than 0.108 moles of PF_5 , XeF_4 is the limiting reactant.

$$0.0909 \text{ mol PF}_5 * \frac{125.97 \text{ g PF}_5}{1 \text{ mol PF}_5} = 11.45 \text{ g PF}_5$$

b. Identify the species being oxidized and the one being reduced.

What is the oxidizing agent, reducing agent?

P is oxidized (+3 to +5), Xe is reduced (+4 to 0). Oxidizing agent: XeF_4 ,
reducing agent: PF_3

Note that oxidizing and reducing agents are the entire compounds themselves, not just the element

9. You combine 0.871 moles of sodium phosphate with 1.23 L of water. What is the molarity of the solution, of sodium ions, and phosphate ions?

$$\text{Molarity of solution} = \frac{0.871 \text{ mol Na}_3\text{PO}_4}{1.23 \text{ L H}_2\text{O}} = 0.708 \text{ M}$$

$$[\text{Na}^+] = \frac{0.871 \text{ mol Na}_3\text{PO}_4}{1.23 \text{ L H}_2\text{O}} * \frac{3 \text{ mol Na}^+}{1 \text{ mol Na}_3\text{PO}_4} = 2.12 \text{ M}$$

$$[\text{PO}_4^{3-}] = \frac{0.871 \text{ mol Na}_3\text{PO}_4}{1.23 \text{ L H}_2\text{O}} * \frac{1 \text{ mol PO}_4^{3-}}{1 \text{ mol Na}_3\text{PO}_4} = 0.708 \text{ M}$$