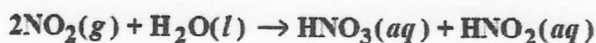
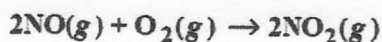
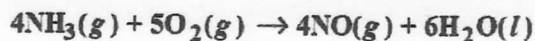


Be sure to answer all parts.

Industrially, nitric acid is produced by the Ostwald process, as represented by the following equations:



What mass of NH_3 (in grams) must be used to produce 3.97 tons of HNO_3 by the Ostwald process, assuming an 80.0 percent yield in each step (1 ton = 2000 lb; 1 lb = 453.6 g)? Enter your answer in scientific notation.

$\times 10^{\text{}$ g NH_3

$$3.97 \text{ tons of } \text{HNO}_3 \times \frac{2000 \text{ lb}}{1 \text{ ton}} \times \frac{453.6 \text{ g}}{1 \text{ lb}} \times \frac{1 \text{ mol } \text{HNO}_3}{63.018 \text{ g } \text{HNO}_3} = 5.715 \times 10^4 \text{ mol } \text{HNO}_3$$

$$5.715 \times 10^4 \text{ mol } \text{HNO}_3 \times \frac{2 \text{ mol } \text{NO}_2}{1 \text{ mol } \text{HNO}_3} \times \frac{2 \text{ mol } \text{NO}}{2 \text{ mol } \text{NO}_2} \times \frac{4 \text{ mol } \text{NH}_3}{4 \text{ mol } \text{NO}} = 1.143 \times 10^5 \text{ mol } \text{NH}_3$$

$$1.143 \times 10^5 \text{ mol } \text{NH}_3 \times \frac{17.034 \text{ g } \text{NH}_3}{1 \text{ mol } \text{NH}_3} = 1.947 \times 10^6 \text{ g } \text{NH}_3$$

if the reactions were 100% efficient, $1.947 \times 10^6 \text{ g } \text{NH}_3$ would be needed to generate 3.97 tons of HNO_3 . Each step yields only 80% of the theoretical amount, however. Since there are 3 steps that are 80% efficient:

$$(0.8)(0.8)(0.8) = 0.512 \quad 51.2\% \text{ total yield}$$

$$x \text{ g of } \text{NH}_3 \text{ Actually needed } (0.512) = 1.95 \times 10^6 \text{ mol } \text{NH}_3 \text{ theoretically needed (1)}$$

$$x \text{ g of } \text{NH}_3 \text{ Actually needed} = \boxed{3.79 \times 10^6 \text{ g } \text{NH}_3}$$

Enter your answer in the provided box.

Assume the atomic mass of element X is 22.99 amu. A 28.88-g sample of X combines with 100.35 g of another element Y to form a compound XY. Calculate the atomic mass of Y.

amu



$$\text{moles of X} = 28.88 \text{ g X} \times \frac{1 \text{ mol X}}{22.99 \text{ g X}} = 1.256 \text{ mol X}$$

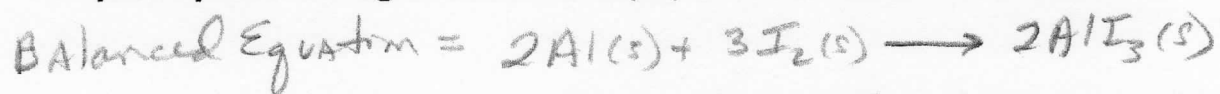
$$1.256 \text{ mol X} \times \frac{1 \text{ mol Y}}{1 \text{ mol X}} = 1.256 \text{ mol Y}$$

Molar mass of Y:

$$\frac{100.35 \text{ g Y}}{1.256 \text{ mol Y}} = \frac{79.90 \text{ g Y}}{\text{mol}} = \text{Atomic mass}$$

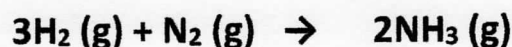
(When completing the problem, you must assume that X reacts completely with Y to generate XY in 100% yield.)

3.50 Calculate the mass in grams of iodine (I₂) that will react completely with 20.4g of aluminum (Al) to form aluminum iodide.



$$20.4\text{g Al} \times \frac{1\text{mol Al}}{26.98\text{g Al}} \times \frac{3\text{mol I}_2}{2\text{mol Al}} \times \frac{253.8\text{g I}_2}{1\text{mol I}_2} = 288\text{g I}_2$$

3.72 Ammonia is prepared by the reaction between hydrogen and nitrogen:



In a particular reaction, 6.0 mol of NH₃ were produced. How many moles of H₂ and how many moles of N₂ were consumed to produce this amount of NH₃?

$$6.0\text{mol NH}_3 \times \frac{3\text{mol H}_2}{2\text{mol NH}_3} = 9.0\text{mol H}_2$$

$$6.0\text{mol NH}_3 \times \frac{1\text{mol N}_2}{2\text{mol NH}_3} = 3.0\text{mol N}_2$$

3.48 The density of water is 1.00g/mL at 4 °C. How many water molecules are present in 15.78 mL of water at this temperature?

$$15.78\text{mL} \times \frac{1.00\text{g}}{\text{mL}} = 15.78\text{g} \times \frac{1\text{mol H}_2\text{O}}{18.02\text{g H}_2\text{O}} \times \frac{6.022 \times 10^{23}\text{ molecules}}{1\text{mol H}_2\text{O}} = 5.275 \times 10^{23}\text{ molecules}$$

3.64 Ascorbic acid contains C, H, and O. In one combustion analysis, 5.24g of ascorbic acid yields 7.86g of CO₂ and 2.14g H₂O. Calculate the empirical formula and molecular formula of ascorbic acid given that its molar mass is about 176g/mol.

Need to know the moles of C, H, + O:

$$7.86g \text{ CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01g \text{ CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 0.1786 \text{ mol C} \times \frac{12.01g \text{ C}}{1 \text{ mol C}} = 2.145g \text{ C}$$

$$2.14g \text{ H}_2\text{O} \times \frac{1 \text{ mol}}{18.02g \text{ H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 0.2375 \text{ mol H} \times \frac{1.008g \text{ H}}{1 \text{ mol H}} = 0.2394g \text{ H}$$

$$5.24g \text{ sample} - (2.145g \text{ C} + 0.2394g \text{ H}) = 2.856g \text{ O} \times \frac{1 \text{ mol O}}{16.00g} = 0.1785 \text{ mol O}$$

$$\text{C} \frac{0.1786}{0.1785} \text{ H} \frac{0.2375}{0.1785} \text{ O} \frac{0.1785}{0.1785} = \text{C}_1 \text{ H}_{1.33} \text{ O}_1 \text{ multiply by 3 to get whole numbers}$$

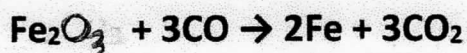
$$\text{Empirical Formula} = \text{C}_3 \text{ H}_4 \text{ O}_3$$

$$\text{Empirical Formula weight} = \underline{\underline{88.06g}}$$

$$\frac{176g}{88.06} = 1.998 \approx 2$$

$$\text{molecular formula} = \text{C}_6 \text{ H}_8 \text{ O}_6$$

3.138 One of the reactions that occurs in a blast furnace, where iron ore is converted to cast iron, is:



Suppose that 1.64×10^3 kg of Fe is obtained from a 2.62×10^3 kg sample of Fe₂O₃. Assuming the reaction goes to completion, what is the percent purity of Fe₂O₃ in the original sample?

The percent yield will be equal to the percent purity of iron III

$$2.62 \times 10^3 \text{ kg Fe}_2\text{O}_3 \times \frac{1000g \text{ Fe}_2\text{O}_3}{1 \text{ kg Fe}_2\text{O}_3} \times \frac{1 \text{ mol Fe}_2\text{O}_3}{159.7g \text{ Fe}_2\text{O}_3} \times \frac{2 \text{ mol Fe}}{1 \text{ mol Fe}_2\text{O}_3} \times \frac{55.85g \text{ Fe}}{1 \text{ mol Fe}} \times \frac{1 \text{ kg Fe}}{1000g \text{ Fe}} = 1.83 \times 10^3 \text{ kg Fe}$$

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{1.64 \times 10^3 \text{ kg Fe}}{1.83 \times 10^3 \text{ kg Fe}} \times 100 = 89.5\% = \text{purity of Fe}_2\text{O}_3$$

3.94 When heated, lithium (Li) reacts with nitrogen (N₂) to form lithium nitride. What is the theoretical yield of lithium nitride in grams when 12.3g of Li is heated with 33.6g of N₂? If the actual yield of lithium nitride is 5.89g, what is the percent yield of the reaction?

① WRITE a balanced equation: $6\text{Li}(s) + \text{N}_2(g) \rightarrow 2\text{Li}_3\text{N}(s)$

$$12.3\text{g Li} \times \frac{1\text{mol Li}}{6.941\text{g Li}} = 1.772\text{mol Li}$$

$$33.6\text{g N}_2 \times \frac{1\text{mol N}_2}{28.02\text{g N}_2} = 1.199\text{mol N}_2$$

$$1.199\text{mol N}_2 \times \frac{6\text{mol Li needed}}{1\text{mol N}_2} = 7.194\text{mol Li needed} \quad \left(\begin{array}{l} \text{Since we don't} \\ \text{have this much} \\ \text{Li is limiting} \end{array} \right)$$

$$1.772\text{mol Li} \times \frac{2\text{mol Li}_3\text{N}}{6\text{mol Li}} \times \frac{34.83\text{g Li}_3\text{N}}{1\text{mol Li}_3\text{N}} = 20.6\text{g Li}_3\text{N} = \text{theoretical yield}$$

$$\text{Since Actual yield} = 5.89\text{g} \Rightarrow \frac{5.89\text{g}}{20.6\text{g}} \times 100 = 28.6\%$$

3.90 Given the reaction:

$\text{CaF}_2 + \text{H}_2\text{SO}_4 \rightarrow \text{Calcium sulfate and hydrofluoric acid}$

Write a balanced equation. In one process, 6.00 kg of CaF₂ is treated with an excess of sulfuric acid and yields 2.86kg of hydrofluoric acid.

Calculate the percent yield of hydrofluoric acid.

$$\text{CaF}_2 + \text{H}_2\text{SO}_4 \rightarrow \text{CaSO}_4 + 2\text{HF}$$

$$6.00 \times 10^3\text{g CaF}_2 \times \frac{1\text{mol CaF}_2}{78.08\text{g CaF}_2} \times \frac{2\text{mol HF}}{1\text{mol CaF}_2} \times \frac{20.008\text{g HF}}{1\text{mol HF}} \times \frac{1\text{kg}}{1000\text{g}} = 3.075\text{kg HF}$$

Since Actual yield is 2.86 kg HF:

$$\frac{2.86\text{kg}}{3.075\text{kg}} \times 100\% = 93.0\%$$

1. Fill in the correct formula for each compound on the line next to its name below:

Uranium (VI) fluoride UF_6
Magnesium hydroxide $\text{Mg}(\text{OH})_2$
Sodium carbonate Na_2CO_3
Potassium sulfite K_2SO_3
Phosphoric Acid H_3PO_4

2. Fill in the correct name for each compound on the line next to its formula below:

$\text{Fe}(\text{OH})_3$ iron (III) Hydroxide
 $\text{Ca}(\text{OCl})_2$ Calcium hypochlorite
 NH_4Cl Ammonium chloride
 $\text{K}_2\text{Cr}_2\text{O}_7$ potassium dichromate
 KH_2PO_4 potassium dihydrogen phosphate

- 2.84 Fill in the correct formula for each compound on the line next to its name below:

Copper (I) cyanide CuCN
Strontium chlorite $\text{Sr}(\text{ClO}_2)_2$
Perbromic acid HBrO_4
Hydroiodic acid HI
Disodium ammonium phosphate $\text{Na}_2(\text{NH}_4)\text{PO}_4$
Potassium dihydrogen phosphate KH_2PO_4
Iodine heptafluoride IF_7
Tetraphosphorous decasulfide P_4S_{10}
Mercury (II) oxide HgO
Cobalt (II) carbonate CoCO_3
Selenium hexafluoride SeF_6
nickel (II) nitrate hexahydrate $\text{Ni}(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O}$

2.82. Fill in the correct name for each compound on the line next to its formula below:

KClO	<u>potassium hypochlorite</u>
Ag ₂ CO ₃	<u>Silver Carbonate</u>
HNO ₂	<u>Nitrous Acid</u>
KMnO ₄	<u>potassium permanganate</u>
CsClO ₃	<u>Cesium chlorate</u>
KNH ₄ SO ₄	<u>potassium Ammonium Sulfate</u>
FeO	<u>iron (II) oxide</u>
Fe ₂ O ₃	<u>iron (III) oxide</u>
TiCl ₄	<u>titanium (IV) chloride</u>
NaH	<u>Sodium hydride</u>
Li ₃ N	<u>Lithium nitride</u>
Na ₂ O	<u>Sodium oxide</u>
Na ₂ O ₂	<u>Sodium peroxide</u>

2.96 What is wrong with the name given for each of the compounds below?

Ba Cl ₂	<u>Barium dichloride</u>	<u>ionic compound - therefore No prefixes</u>
Fe ₂ O ₃	<u>iron (II) oxide</u>	<u>Iron has +3 charge - therefore iron(III) oxide</u>
CsNO ₂	<u>cesium nitrate</u>	<u>NO₂⁻ is nitrite ion - therefore cesium nitrite</u>
Mg(HCO ₃) ₂	<u>magnesium (II) bicarbonate</u>	

Not Needed since Mg ion is always Mg⁺²
 since Magnesium is in second group of
 periodic table: Alkaline Earth metal,