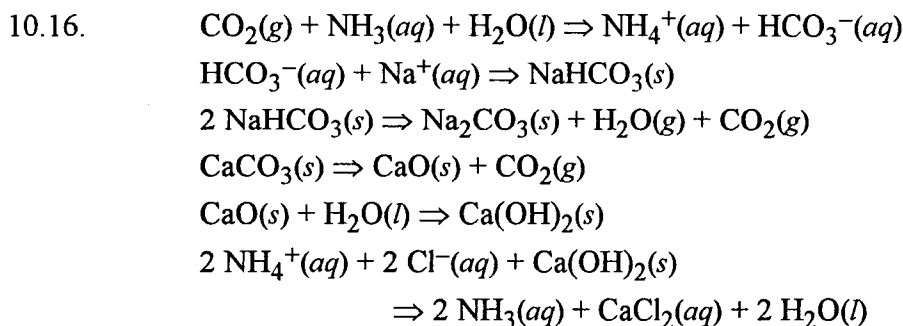


## Chapter 10

### THE GROUP 1 ELEMENTS: THE ALKALI METALS

#### Exercises

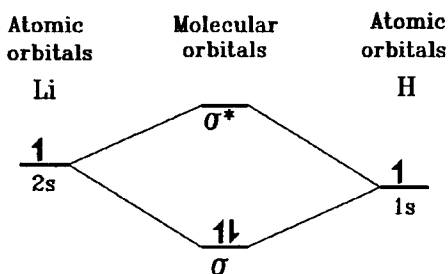
- 10.2. (a)  $6 \text{Li}(s) + \text{N}_2(g) \Rightarrow 2 \text{Li}_3\text{N}(s)$   
(b)  $2 \text{CsO}_2(s) + 2 \text{H}_2\text{O}(l) \Rightarrow 2 \text{CsOH}(aq) + \text{H}_2\text{O}_2(aq) + \text{O}_2(g)$   
(c)  $2 \text{NaHCO}_3(s) \Rightarrow \text{Na}_2\text{CO}_3(s) + \text{CO}_2(g) + \text{H}_2\text{O}(g)$   
(d)  $\text{NH}_4\text{NO}_3(s) \Rightarrow \text{N}_2\text{O}(g) + 2 \text{H}_2\text{O}(g)$
- 10.4. Lithium. This metal has the most negative standard reduction potential; that is, it is the most strongly reducing of any element. Its low chemical reactivity must indicate a higher activation energy of reaction than that of the other alkali metals and hence a slower reaction.
- 10.6. Sodium. Hydration is favored by high charge density cations; and of the two monovalent ions, sodium is the smaller and will have the higher charge density.
- 10.8. The greater the difference in size, the lower the lattice energy, and, as this is an energy input term in the solution process, the solubility will be higher. The lithium salt would give the smallest alkali metal cation to pair with the large organic anion and hence the greatest solubility.
- 10.10. Because the equilibrium of the synthesis reaction  
$$\text{Na}(l) + \text{KCl}(l) \Rightarrow \text{K}(l) + \text{NaCl}(l)$$
lies to the left. To make the reaction shift right, the potassium, which boils at a lower temperature than sodium, must be continuously removed as a gas.
- 10.12. (a) Sodium hydroxide; (b) anhydrous sodium carbonate; (c) sodium carbonate decahydrate.
- 10.14. (a) Loss of water by a hydrated salt in a low humidity environment.  
(b) Chemical similarities of one element and the element to its lower right in the periodic table.



The problems are the disposal of waste calcium chloride and the high energy requirements of the process.

10.18. The ammonium ion is monovalent, like the alkali metals (unlike most other metals that have 2+ or higher charge); its salts are all soluble like those of the alkali metal salts; its size is about the middle of the alkali metal ion range; all its common salts are colorless, like those of the alkali metals.

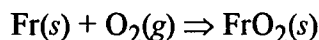
10.20. There will be a single covalent bond.



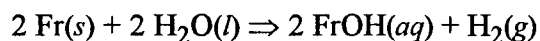
10.22. Potassium ion is inside and sodium ion outside of cells.

### Beyond the Basics

10.24. Francium would be predicted to be a liquid at room temperature; the most reactive of the alkali metals; the most ionic of the series (that is, the most electropositive element); the largest atom and cation of the series. It should rapidly form a dioxide (1-) on exposure to air —



— and it should react explosively with water:

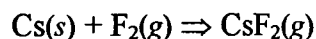
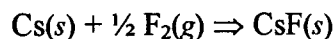


The large cation should form the least soluble salts of the alkali metals with large anions. The salts of the very low charge density francium are unlikely to be hydrated. Other reactions common to the alkaline earth metals should be expected.

10.26. The  $\text{PtF}_6^-$  ion must be a very large, low charge anion. Lithium is the smallest of the alkali metal cations. It will not stabilize the large anion; instead it abstracts the fluoride ion. Cesium or any of the other lower alkali metals should form stable compounds, such as  $\text{CsPtF}_6$ .

10.28. Lithium consists of two isotopes, lithium-6 (7 percent) and lithium-7 (93 percent). The higher average atomic mass must be due to the sample being depleted of lithium-6. Lithium-6 is used in nuclear weapons and the remaining lithium, enriched in lithium-7, is sold to chemical suppliers. Samples of commercial lithium compounds have been analyzed that contain as little as 3.75 percent lithium-6.

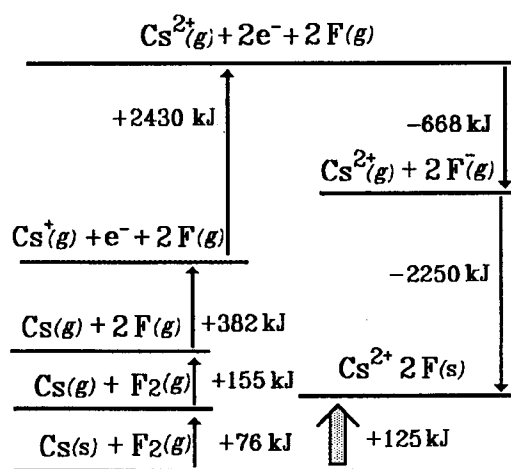
10.30. We can solve the problem once we know the enthalpies of formation of the two compounds



Reversing the second equation and summing gives



The enthalpy of formation of  $\text{CsF}$  is  $-554 \text{ kJ}\cdot\text{mol}^{-1}$  while that for  $\text{CsF}_2$  can be found from the Born-Haber cycle as  $+125 \text{ kJ}\cdot\text{mol}^{-1}$ :



Thus the enthalpy change for the decomposition will be

$$[(-554) - (+125)] \text{ kJ}\cdot\text{mol}^{-1} = -679 \text{ kJ}\cdot\text{mol}^{-1}$$

As  $\frac{1}{2}$  mol of gas is produced in the decomposition, the entropy change will be positive.

As  $\Delta G = \Delta H - T\Delta S$ ,  $\Delta G$  will be negative and the decomposition thermodynamically spontaneous.