

Chapter 11

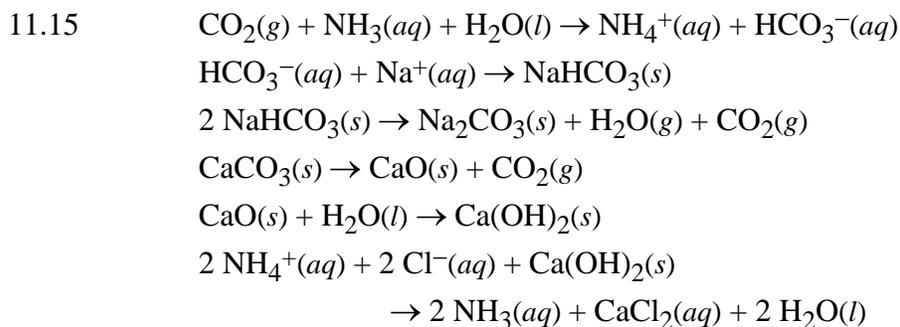
THE GROUP 1 ELEMENTS: THE ALKALI METALS

Exercises

- 11.1 (a) $2 \text{Na}(s) + 2 \text{H}_2\text{O}(l) \rightarrow 2 \text{NaOH}(aq) + \text{H}_2(g)$
(b) $\text{Rb}(s) + \text{O}_2(g) \rightarrow \text{RbO}_2(s)$
(c) $2 \text{KOH}(s) + \text{CO}_2(g) \rightarrow \text{K}_2\text{CO}_3(s) + \text{H}_2\text{O}(l)$
(d) $2 \text{NaNO}_3(s) \rightarrow 2 \text{NaNO}_2(s) + \text{O}_2(g)$
- 11.3 They resemble “typical” metals in that they are shiny and silvery and good conductors of heat and electricity. The alkali metals differ from “typical” metals in that they are soft, extremely chemically reactive, and have low melting points and very low densities.
- 11.5 Any three of the following:
All common chemical compounds are water soluble.
They always form ions of +1 oxidation state.
Their compounds are almost always ionic.
The low-charge-density alkali metal ions stabilize large low-charge anions such as hydrogen carbonate.
Their compounds are rarely hydrated.
- 11.7 The most likely argument is that the hydroxide ion can hydrogen bond with the surrounding water molecules whereas the chloride ion can only form ion-dipole attractions. This hydrogen bond formation would release more energy than the ion-dipole formation. Alternatively, we can simply argue that the smaller hydroxide ion will have a higher charge density and the ion-dipole attractions will be greater—this is the same argument, just a different approach. In fact, the hydration enthalpy for the hydroxide ion is indeed much greater than that of the chloride ion.
- 11.9 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.

- 11.11 (a) Sodium hydroxide; (b) anhydrous sodium carbonate; (c) sodium carbonate decahydrate.
- 11.13 (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.

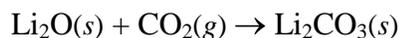
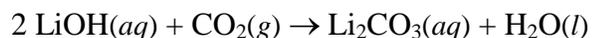
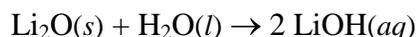
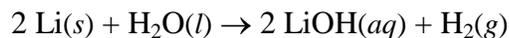
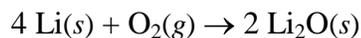
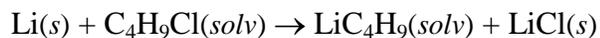
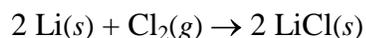
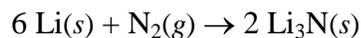


The problems are the disposal of waste calcium chloride and the high energy requirements of the process.

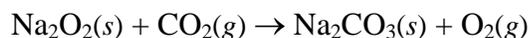
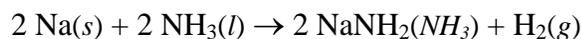
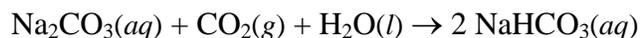
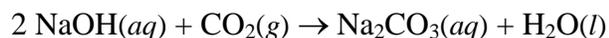
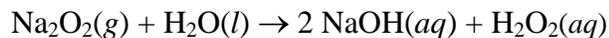
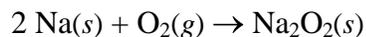
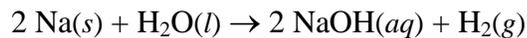
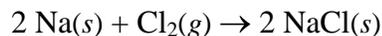
- 11.17 The ammonium ion is monovalent, like the alkali metals (unlike most other metals that have 2+ or higher charge); its salts are all soluble, like those of the alkali metal salts; its size is about the middle of the alkali metal ion range; all its common salts are colorless, like those of the alkali metals.
- 11.19 Potassium dioxide(1-) has a lower molar mass than the equivalent cesium compound. For launch, minimum mass for the same oxygen-generating capacity is crucial. Also, the common potassium salts are much lower in cost than those of cesium.

11.21 The ammonium ion is large and hence resembles the lower alkali metal ions. In particular, its non-spherical shape makes it resemble the largest of the alkali metal ions.

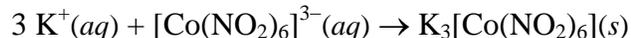
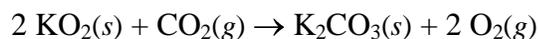
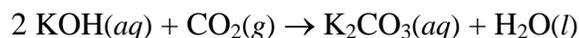
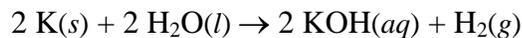
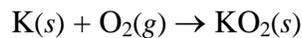
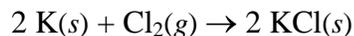
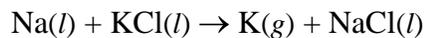
11.23 Lithium:



Sodium:



Potassium



Beyond the Basics

$$11.25 \text{ Mass of sodium} = 1.00 \text{ t} \times (10^6 \text{ g}) / (1 \text{ t}) = 1.00 \times 10^6 \text{ g}$$

$$\text{Mole sodium} = (1.00 \times 10^6 \text{ g}) \times (1 \text{ mol}) / (23.0 \text{ g}) = 4.35 \times 10^4 \text{ mol}$$



$$\text{Thus mole electrons} = 4.35 \times 10^4 \text{ mol}$$

$$\text{Current} = (4.35 \times 10^4 \text{ mol}) \times \frac{(9.65 \times 10^4 \text{ A} \cdot \text{V} \cdot \text{s})}{(1 \text{ mol})(7.0 \text{ V})(86400 \text{ s})} = 6.94 \times 10^4 \text{ A}$$

11.27 In the series LiF to CsF, there is an increasing mismatch in ion sizes; thus the lattice energy will decrease more than otherwise expected. As a result, the enthalpies of formation will decrease. For the series LiI to CsI, there is a decreasing mismatch in ion sizes; thus the lattice energy will decrease to a lesser extent than otherwise expected. As a result the enthalpies of formation will increase.

11.29 Sodium fluoride will be the less soluble because there is a close match in ion sizes resulting in a higher lattice energy ($\text{Na}^+ = 116 \text{ pm}$; $\text{F}^- = 117 \text{ pm}$). There is a mismatch in sizes with the tetrafluoroborate ion (which actually has a radius of 218 pm). Thus the hydration energy will more probably exceed the (lower) lattice energy, making the compound more soluble.

11.31 There are two possible answers: that there is appreciable covalent bonding in the lithium hydride, thus reducing the Li–H separation, or that the lithium ion is so small that the lattice consists of touching hydride ions with lithium ions “rattling around” in the lattice holes—this is certainly true of lithium iodide.

11.33 LiF and KI.

11.35 Calcium-40 is a “doubly magic” nucleus with filled shells of protons and neutrons.