Chapter 10

HYDROGEN

Exercises

10.2 (a) A solid species that forms an open lattice structure having another species locked within the interstices without any chemical bonding occurring.
(b) A two-dimensional diagram that displays the variation of phase of an element or compound with changes in pressure and temperature.

10.4 Only oxygen-17, as both carbon-12 and oxygen-16 have even numbers of protons and neutrons.

10.6 Unlike the alkali metals, hydrogen is not a metal, nor does it react with water.

10.8 Dihydrogen is a comparatively unreactive gas because it has a very high bond energy and which is stronger than the bond of hydrogen to most other elements.

10.10 (a) $WO_3(s) + 3\ H_2(g) \rightarrow W(s) + 3\ H_2O(g)$

(b) $H_2(g) + Cl_2(g) \rightarrow 2\ HCl(g)$

(c) $2\ Al(s) + 6\ HCl(aq) \rightarrow 2\ AlCl_3(aq) + 3\ H_2(g)$

10.12 For the reaction,

$\Delta H^\circ = [1(-394) + 2(-286) - 1(-75)]\ kJ \cdot mol^{-1}$

$= -891\ kJ \cdot mol^{-1}$

$\Delta S^\circ = [1(+214) + 2(+70) - 1(+186) - 2(+205)]\ J \cdot mol^{-1} \cdot K^{-1}$

$= -242\ J \cdot mol^{-1} \cdot K^{-1} = -0.242\ kJ \cdot mol^{-1} \cdot K^{-1}$

$\Delta G^\circ = (-891\ kJ \cdot mol^{-1}) - (298\ K)(-0.242\ kJ \cdot mol^{-1} \cdot K^{-1})$

$= -819\ kJ \cdot mol^{-1}$

10.14 Ionic hydrides are solids; covalent hydrides are mostly gases or low-boiling liquids.

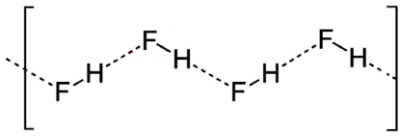
10.16 (a) Metallic; (b) none; (c) covalent; (d) ionic.

10.18 Using Allred-Rochow electronegativities:

	CH_4	SiH_4	SnH_4
ΔH_f^o	–75	+34	+188
E.D.	–0.30	+0.46	+0.48

Thus the only thermodynamically stable hydride (methane) has a bond polarity in which the hydrogen is partially positive. The other two members of the group, having a partially negative hydrogen, are thermodynamically unstable. There is a rough trend of increasing free energy with bond polarity.

10.20 A hydrogen fluoride molecule can only form two hydrogen bonds:



A water molecule can form four hydrogen bonds (each oxygen atom can hydrogen-bond to two hydrogen atoms). Thus although each hydrogen bond in water is weaker (lower electronegativity difference) the existence of twice as many increases the attractions between neighboring water molecules and raises the melting point.

10.22 $2\ Na(s) + H_2(g) \rightarrow 2\ NaH(s)$

$H_2(g) + F_2(g) \rightarrow 2\ HF(g)$

$2\ Ti(s) + 1.9\ H_2(g) \rightarrow 2\ TiH_{1.9}(s)$

$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$

$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$

$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(l)$

Beyond the Basics

10.24 (a) $B_2H_6(g) + 3\ O_2(g) \rightarrow B_2O_3(s) + 3\ H_2O(l)$

(b) $PbH_4(g) + 2\ O_2(g) \rightarrow PbO_2(s) + 2\ H_2O(l)$

(c) $2\ BiH_3(g) + 3\ O_2(g) \rightarrow Bi_2O_3(s) + 3\ H_2O(l)$

10.26 (a) In interstitial hydrides, the hydrogen atoms are occupying spaces between the metal atoms. This can result in the lattice swelling and hence a lower density.

(b) In ionic hydrides, the metals are present as the cation. The ionic radius of a metal is always much smaller than that of the metallic radius. For ionic hydrides, it is sometimes better to consider that the lattice consists of the hydride ions with the metal cations occupying the interstices. As a result, the density of the ionic solid could be higher than that of the metal itself.