10.12 temperature and # of moles stays constant while pressure + volume are changed, so use Boyle's Law:

$$P_1V_1 = P_2V_2 \quad \text{ǂ} \quad V_2 = \frac{P_1V_1}{P_2}$$

so, $V_2 = \frac{(760 \text{ torr})(22.4 L)}{616 \text{ torr}} = 27.6 L$

10.16 (a) $t_F = \left[\frac{9}{5}(342.45) + 32\right]°F = 648.41°F$

(b) $t_C = 5/9(98.6 - 32)°C = 37.0°C$ (body temperature)

(c) $t_F = \left[\frac{9}{5}(-10.6) + 32\right]°F = 12.9°F$

(d) $t_C = 5/9(-40.0 - 32)°C = -40.0°C$ (the two scales coincide here)

2.19 # of moles and pressure remains constant while temperature and volume are changed, so use Charle's Law:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \text{ǂ} \quad V_1 = \frac{V_2T_1}{T_2}$$

so, $V_1 = V_2 \frac{T_1}{T_2} = \frac{2.38L \times 673K}{773K} = 2.07L$

2.22 Find the # of moles then use the conversion:

$$1 \text{ mole} = 6.02 \times 10^{23} \text{ atoms}$$

(a) $1 \text{ mole} = 6.02 \times 10^{23}$ atoms

(b) $PV = nRT$ so $n = PV/RT = \frac{(2.00 \text{ atm})(22.4L)}{(0.08206 \text{ atm L mol}^{-1} K^{-1})(273K)} = 2.00$ mole

$2.00 \text{ mole} \times \frac{6.02 \times 10^{23} \text{ atoms}}{1 \text{ mole}} = 1.20 \times 10^{24}$ atoms

(c) as above, $n = 0.500 \text{ mol} \times \frac{6.02 \times 10^{23} \text{ atoms}}{1 \text{ mole}} = 3.01 \times 10^{23}$ atoms

(d) as above, $n = 0.500 \text{ mol} \times \frac{6.02 \times 10^{23} \text{ atom}}{1 \text{ mol}} = 3.01 \times 10^{23}$ atom

(e) as above, $n = 2.00$ mole so # atoms $= 1.20 \times 10^{24}$ atoms